11 Calorimetry

Introduction

Heat is the energy transferred between systems at different temperatures. Heat only describes energy transferred, and hence, we do not speak of objects as containing heat. If heat flows from a table top into a cold glass of water sitting on it, the internal energy of the glass of water increases and the internal energy of the table top decreases by the same amount. However, the temperatures of the two systems does not in general change by the same amount. It is important *not* to equate temperature with internal energy. Different materials have different relationships between internal energy and temperature.

The **specific heat** relates changes in internal energy to changes in temperature. The specific heat c of a material is the amount of energy in Joules required to raise the temperature of one kilogram of a given material by one Celsius degree. A quantity of heat Q flowing into an object of mass m and specific heat c is related to the resulting change in temperature by

$$Q = mc\Delta T. \tag{1}$$

Changes of phase involve the absorption or release of energy. For example, it costs a certain amount of energy to break the bonds between the molecules in an ice crystal in order to turn it into liquid water. The amount of energy in Joules released or absorbed per kilogram of material in a particular change of phase is called the **heat of transformation** L of that phase change. The heat of transformation associated with the change from solid to liquid phase is known as the **heat of fusion**, and the heat of transformation associated with the change from liquid to vapor phase is known as the **heat of vaporization**.

The heat flowing into or out of an object of mass m as a result of a phase change with heat of transformation L is given by

$$Q = mL. \tag{2}$$

Phase changes happen at fixed temperatures, so there is no temperature change in the expression for the heat of transformation. In changes from solid to liquid or from liquid to vapor phases, heat flows into the system. In the reverse of these transitions, heat flows out of the system.

The law of conservation of energy dictates that there is no net heat flow within a closed system,

$$Q_A + Q_B + Q_C ... = 0 \tag{3}$$

For example, if I submerge an ice cube in water in a Styrofoam cup and allow them to come to the same temperature, the heat flowing out of the water must equal the heat flowing into the ice cube to the extent that no heat is gained or lost to the surrounding environment.

You will apply the law of conservation of energy to an approximately closed system – a calorimeter. You will determine the specific heats of samples of three common elements by heating each sample and combining it with cold water in a calorimeter and measuring the changes in temperature of the calorimeter, water, and sample. You will also determine the densities of the samples. Then, you will try to identify the samples based on your results.

Experiments and Analysis

Logger Pro Setup

You have been provided with an alcohol thermometer and two temperature probes. Check the temperature probe calibrations as follows.

1. Get a cup of ice water and a cup of hot ($> 50°C$) water from the front of the lab.

2. Connect the two temperature probes to channels 1 and 2 of the LabPro interface (labeled `CH1` and `CH2`).

3. Run `Logger Pro`. The program should automatically recognize the temperature sensors and set up a graph of the two temperatures vs. time.

4. Place both temperature probes and the alcohol thermometer in the hot water and press the `Collect` button 

. Stir the water with the probes for about 10 seconds until the temperature readings reach a constant value, and compare the temperature probe readings with the temperature indicated by the alcohol thermometer.

5. Repeat this process with the ice water.

6. If there is significant disagreement, ask your instructor or TA for help.

Mystery Samples

You have been provided with three cylindrical samples of different materials suggestively labeled C, L, and Z.

1. Make whatever measurements you need to make in order to determine the densities of the samples, and the associated uncertainties, as precisely as you can. (Note that the hooks can be removed from the samples.)

2. Measure the mass of the dry calorimeter cup.

3. Place about 100 mL of cold water and one of the temperature probes in the calorimeter cup, and install the calorimeter cup in its housing.

4. Collect some hot ($T > 50°$ C) water from the container at the front of the lab in a plastic cup. Place one of the three samples in the hot water along with the other temperature probe.

5. Reduce the temperature of the water and calorimeter cup to about 5°C with the help of some ice. **Remove any solid ice from the water before proceeding.**

6. Start a long (5 minutes is plenty) measurement with `Logger Pro`.

7. Collect data for at least 10 seconds to establish the starting temperatures of the hot water and the calorimeter.

8. Using the piece of string tied to the hot sample, transfer it to the calorimeter. **Leave the hot temperature probe in the hot water.**

   To minimize energy losses to the environment, avoid touching the sample, and transfer it to the calorimeter as quickly as possible.

9. Stir the water in the calorimeter with the cold temperature probe, and continue collecting data until the system reaches a stable equilibrium temperature, and then save your measurement.

10. Measure the mass of the water and calorimeter cup so that you can determine the mass of water you used.

11. Repeat this process with the other samples.

12. Make several repeated measurements of the sample that causes the largest change in temperature of the calorimeter.

13. When you pull a sample out of the hot water, it carries some hot water with it, mainly on top. Measure the mass of a wet sample to determine how much water it carries. (You'll need this for the individual assignment.)

Analysis

The law of conservation of energy (Eq. 3) for the calorimeter cup, the water, and the sample is given by

$$m_{\text{water}}\, c_{\text{water}}\, \Delta T_{\text{water}} + m_{\text{cup}}\, c_{\text{cup}}\, \Delta T_{\text{cup}} + m_{\text{sample}}\, c_{\text{sample}}\, \Delta T_{\text{sample}} = 0 \tag{4}$$

You have measured the masses as well as the initial and final temperatures of all three objects. The calorimeter cup is aluminum with specific heat $c_{\text{Al}} = 900\ \frac{\text{J}}{\text{kg}^\circ\text{C}}$. The specific heat of water is $c_{\text{water}} = 4190\ \frac{\text{J}}{\text{kg}^\circ\text{C}}$.

1. Solve Eq. 4 for c_{sample} and set up a spreadsheet to calculate the densities and specific heats of the samples from your measurements.

2. Use the repeated measurements you made of one sample to determine the average value and the standard deviation of the mean (SDOM).

3. Use the ratio of the standard deviation of your repeated measurements (*not* the SDOM) to the average value to assign statistical uncertainties to your specific heat results for the other samples.

4. Each sample is made of a single element. Based on your density and specific heat results, determine which elements are compatible, within uncertainty, with your results for each sample. (Refer to Appendix F of your text for densities and specific heats.)

Before You Leave Lab

Show your work to your lab instructor and give preliminary answers to questions

Group Assignment

Hand in your spreadsheet and answers to the following questions.

1. What elements are compatible with your density and specific heat measurements of each sample? Explain.

2. Why was it important to remove any solid ice from the calorimeter in step 5 of the experiment?

3. When you transferred the samples from the hot water into the calorimeter, some hot water rode in on the sample. Estimate how large a difference ignoring this small amount of hot water made in your result for the specific heat of one of your samples. Show your work.

4. Why did you use the standard deviation and not the SDOM of your repeated measurements to assign uncertainties to the other samples in step 3 of the analysis above?